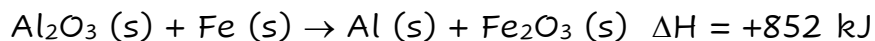


## Enthalpy and Stoichiometry

1. Consider the following unbalanced equation. How many kJ of heat are consumed when 5.69 moles of solid iron react? Is this reaction endothermic or exothermic?



The reaction is **endothermic**



$$5.69 \text{ mol Fe} \times \frac{852 \text{ kJ}}{2 \text{ mol Fe}} = 2.42 \times 10^3 \text{ kJ}$$

2. Consider the following reaction:



- a) is the reaction endothermic or exothermic?

**endothermic**

- b) Calculate the enthalpy change when 32.65 g of NO is produced.

$$M_m \text{ NO} = 30.01 \text{ g/mol}$$

$$32.65 \text{ g NO} \times \frac{1 \text{ mol NO}}{30.01 \text{ g NO}} \times \frac{181.8 \text{ kJ}}{2 \text{ mol NO}} = 98.90 \text{ kJ}$$

3. Write a balanced chemical equation for the combustion of methanol, CH<sub>3</sub>OH in oxygen (O<sub>2</sub>); the reaction gives off 727 kJ of heat. Calculate the enthalpy change when 12.56 g of methanol undergoes combustion in excess oxygen.



$$M_m \text{ CH}_3\text{OH} = 32.04 \text{ g/mol}$$

$$12.56 \text{ g CH}_3\text{OH} \times \frac{1 \text{ mol CH}_3\text{OH}}{32.04 \text{ g}} \times \frac{727 \text{ kJ}}{2 \text{ mol CH}_3\text{OH}} = 142 \text{ kJ}$$

4. The enthalpy change when 1 mole of CH<sub>4</sub> is burned is -890 kJ. To vaporize one mole of water it takes 44.0 kJ of heat. What mass of methane must be burned (in oxygen) to provide the heat required to vaporize 1.50 g of water?



$$1.50 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g}} \times \frac{44.0 \text{ kJ}}{1 \text{ mol H}_2\text{O}} \times \frac{1 \text{ mol CH}_4}{890 \text{ kJ}} \times \frac{16.04 \text{ g CH}_4}{1 \text{ mol CH}_4} = 0.0660 \text{ g CH}_4$$